AP® Chemistry
2002 Sample Student Responses
Form B

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3. Nitrogen monoxide, $NO(g)$, and carbon monoxide, $CO(g)$, are air pollutants generated by automobiles. It has been proposed that under suitable conditions these two gases could react to form $N_2(g)$ and $CO_2(g)$, which are components of unpolluted air.

(a) Write a balanced equation for the reaction described above. Indicate whether the carbon in CO is oxidized or whether it is reduced in the reaction. Justify your answer.

(b) Write the expression for the equilibrium constant, K_p, for the reaction.

(c) Consider the following thermodynamic data.

	NO	CO	CO_2
ΔG_f° (kJ mol^{-1})	+86.55	−137.15	−394.36

(i) Calculate the value of ΔG° for the reaction at 298 K.

(ii) Given that ΔH° for the reaction at 298 K is -746 kJ per mole of $N_2(g)$ formed, calculate the value of ΔS° for the reaction at 298 K. Include units with your answer.

(d) For the reaction at 298 K, the value of K_p is 3.3×10^{120}. In an urban area, typical pressures of the gases in the reaction are $P_{NO} = 5.0 \times 10^{-7}$ atm, $P_{CO} = 5.0 \times 10^{-5}$ atm, $P_{N_2} = 0.781$ atm, and $P_{CO_2} = 3.1 \times 10^{-4}$ atm.

(i) Calculate the value of ΔG for the reaction at 298 K when the gases are at the partial pressures given above.

(ii) In which direction (to the right or to the left) will the reaction be spontaneous at 298 K with these partial pressures? Explain.

a) $2NO(g) + 2CO(g) \longrightarrow N_2(g) + 2CO_2(g)$

b) $K_p = \dfrac{[P_{CO_2}]^2[P_{N_2}]}{[P_{CO}]^2[P_{NO}]^2}$

c) i) $\Delta G^\circ = n(\Delta G_f^\circ)_{products} - n(\Delta G_f^\circ)_{reactants}$

$= (2 \times -394.36) - (2 \times -137.15 + 2 \times 86.55)$

$= -687.52$ KJ

ii) $\Delta G^\circ = \Delta H^\circ - T\Delta S^\circ \Rightarrow \Delta S^\circ = \dfrac{\Delta H^\circ - \Delta G^\circ}{T} = \dfrac{-746\text{KJ} + 687.52\text{KJ}}{298\text{K}} = \dfrac{-58.48\text{KJ}}{298\text{K}}$

$= -0.196$ KJ/K $= -196$ J/K

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ADDITIONAL PAGE FOR ANSWERING QUESTION 3.

d) i) $\Delta G = \Delta G^\circ + RT \ln K_p$

$$K_p = \left[\frac{0.781(3.1\times10^{-4})^2}{(5\times10^{-5})^2(5\times10^{-7})^2}\right] = 1.20\times10^{14}$$

$$= -687.52\,kJ + (8.31\times10^{-3})(298)\ln 1.20\times10^{14}$$

$$= -607.24\,kJ$$

ii) To the left, because ΔG is negative at 208K; this means that reaction will favour moving to the left. The reaction is considered spontaneous if ΔG is -ve. Spontaneous if it produces the products.

3B,

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a) $2NO(g) + 2CO(g) \rightarrow N_2(g) + 2CO_2(g)$

The carbon is oxidized because in the reactant, CO, is neutral however in the product CO_2 is also neutral even with the addition of another oxygen molecule. Therefore, the carbon must have lost a couple of electrons for the oxygen molecule to be attached.

b) $K_p = \frac{[CO_2]^2}{[NO]^2[CO]^2}$

c)(i) $\Delta G^\circ = \sum G^\circ_{product} - \sum G^\circ_{reactant}$

$= (-394.36\text{ kJ}) - (2 \cdot 86.55\text{ kJ} + 2 \cdot -137.15\text{ kJ})$

$= -687.52\text{ kJ}$

(ii) $\Delta G^\circ = \Delta H^\circ - T\Delta S^\circ$

$\Delta S^\circ = -\frac{\Delta G^\circ - \Delta H^\circ}{T}$

$= \frac{-687.52\text{ kJ} - (-746\text{ kJ})}{298\text{ K}}$

$= 0.196\text{ kJ/mol·K}$

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ADDITIONAL PAGE FOR ANSWERING QUESTION 3.

d) (i) $\Delta G^\circ = -RT \ln K_p$

$= -(0.0821 \frac{atm}{n \cdot k})(298) \cdot \ln(3.3 \times 10^{120})$

$= -24.5 \ln(3.3 \times 10^{120})$

(ii) To the left because since ΔG° is negative in the forward (right) direction, meaning its not spontaneous, the left direction would mean spontaneous.